Student Textbook page 636

Q1. Macroscopic properties are observable properties such as temperature, colour, and pH. Microscopic properties are only present at the molecular level.

Q2. Equilibrium can only be established in a closed system. Equilibrium is established when forward and reverse reactions are taking place at the same rate.

Q3. During a dynamic equilibrium, changes are still occurring at the microscopic level. Individual molecules continue to react even though the system displays constant macroscopic properties.

Student Textbook page 638

1. The physical changes most likely to be mentioned involve changes of state. Answers should indicate an open system, probably involving energy changes. Examples include boiling water in an open saucepan on a stove, or melting a block of ice. Another common answer would involve dissolving a solute in water to make an unsaturated solution. Chemical processes that are examples of reversible changes not at equilibrium include the electrolysis of water and the voltaic cell reactions students learned in the unit on electrochemistry.

2. Investigation 16.A modelled a reversible change in a closed system that had constant macroscopic properties at equilibrium. The fourth condition, that equilibrium can be approached from either direction was not modelled. Students should suggest simple modifications to the activity, such as adding water to the product cylinder, or removing water from either one of the cylinders. If time permits, they could repeat the exercise making the changes they have suggested. They should repeat the calculation for the ratio $\frac{\text{volume of product}}{\text{volume of reactant}}$ to make sure it has not changed, within experimental errors.

3. The rate of a reaction depends on the concentration of reactants and the rate constant. If the reaction is started with only reactants present, the rate of the forward reaction is relatively large while the rate of the reverse reaction is essentially zero. As reactants form products, the concentration of reactants decreases and the rate of the forward reaction also decreases. As the concentration of products increases, the rate of the reverse reaction also increases. At equilibrium, both forward and reverse reactions are taking place at the same rate. At the molecular level, particles are constantly colliding and taking part in reactions, even though there is no overall change in the reaction mixture. This is the nature of a dynamic equilibrium.

4. (a) The sealed drink bottle contains an aqueous solution of carbon dioxide in equilibrium with $CO_2(g)$. The concentration of $CO_2(g)$ above the liquid is greater than the concentration of $CO_2(aq)$. Some $CO_2(g)$ molecules will strike the aqueous solution with sufficient energy to be dissolved. At the same time, some $CO_2(aq)$ molecules will have enough energy to leave the solution and enter the space above the liquid as a gas. On average, the number of $CO_2(g)$ molecules entering the solution in a given time will equal the number of $CO_2(aq)$ molecules leaving the solution.

(b) The system includes the solution and the gas above it. The following properties will be constant: temperature; pH of the solution; concentration of dissolved carbon dioxide; appearance of the solution; pressure of $CO_2(g)$ above the liquid.

5. Ice and water form an equilibrium mixture (slush) at 0 °C and 1 atmosphere pressure. At temperatures above 0 °C, the ice will melt. At temperatures below 0 °C, the water will freeze.

6. (a) The system has reached equilibrium. The forward reaction has formed NOBr(g), and the decomposition of NOBr(g) produces NO(g) and $Br_2(g)$. Because the forward reaction rate is the same as the reverse reaction rate, a mixture of NO(g), $Br_2(g)$ and NOBr(g) will be present in the flask.

(b) This is a homogeneous equilibrium because all the substances taking part in the reactions are gaseous.

7. (a) $PCl_5(g) \rightleftharpoons PCl_3(g) + Cl_2(g)$ homogeneous equilibrium

(b) $CaCO_3(s) \rightleftharpoons CaO(s) + CO_2(g)$ heterogeneous equilibrium

(c) $HF(aq) + H_2O(\ell) \rightleftharpoons H_3O^+(aq) + F^-(aq)$ heterogeneous equilibrium

(d) $N_2O_3(g) \rightleftharpoons NO_2(g) + NO(g)$ homogeneous equilibrium

(e) $C(s) + CO_2(g) \rightleftharpoons 2CO(g)$ heterogeneous equilibrium

8. The double arrow, $\rightleftharpoons$, in a chemical equation indicates that an equilibrium mixture will be formed containing quantities of reactants and products.

9. False. Equilibrium occurs when rates of opposing changes are equal.

10. (a) The towel in the gym bag remains damp because it is in a relatively closed system. There is not enough moisture leaving the system to dry the towel.

(b) A running clothes dryer that is vented to the outside is an open system. The vent gives water vapour a way out.

11. False. Adding heat to a system at equilibrium will change equilibrium.

Student Textbook page 641

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

1. $K_c = \dfrac{[CH_3COOCH_2CH_3][H_2O]}{[CH_3CH_2OH][CH_3COOH]}$

2. $K_c = \dfrac{[NO]^2}{[N_2][O_2]}$

3. $K_c = \dfrac{[H_2O]^2}{[H_2]^2[O_2]}$

4. $K_c = \dfrac{[Fe^{2+}]^2[I_2]}{[Fe^{3+}]^2[I^-]^2}$

5. $K_c = \dfrac{[NO]^4[H_2O]^6}{[NH_3]^4[O_2]^5}$

Student Textbook page 644

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

6. $K_c = 1.9 \times 10^{-2}$

7. $K_c = 1.2 \times 10^{2}$

8. $K_c = 0.013$

9. $[S_2(g)] = 0.046$ mol/L

10. $[CH_4(g)] = 0.11$ mol/L

Student Textbook pages 649–650

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

11. The value of K_c is smaller at 1100 K. The position of equilibrium lies farther to the left and favours the formation of $H_2(g)$ and $I_2(g)$ at the higher temperature of 1100 K.

12. The order favouring the formation of products is reaction (III) > (II) > (I).

13. **(a)** proceeds to completion

(b) essentially does not proceed to right

(c) proceeds to completion

14. There is essentially no dissociation of $HCl(g)$ into $H_2(g)$ and $Cl_2(g)$.

15. $[Ag(NH_3)_2]^+(aq)$ is more stable

Student Textbook page 651

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

16. The reaction shifts to the left.

17. **(a)** As the reaction proceeds there is no change in the number of gas molecules on each side of the equation. Therefore, increasing the volume of the container has no effect on the position of equilibrium.

(b) There are more gas molecules on the left side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the left.

(c) There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.

(d) There are more gas molecules on the right side of the equation. Therefore, increasing the volume of the container causes the reaction to shift to the right.

18. (a) left

(b) right

(c) no change

(d) left

(e) no change

19. Change **(d)** affects the value.

20. If the reactant, methylcyclohexane, $C_7H_{14}(g)$, is added, the equilibrium will shift to the right because this will use up this reactant and minimize the change. Therefore, increasing $[C_7H_{14}(g)]$ increases the production of toluene. If either product, toluene ($C_7H_8(g)$) or hydrogen ($H_2(g)$), is removed, the equilibrium will shift to the right because this will produce more of these products and minimize the change. Therefore, decreasing either $[C_7H_8(g)]$ or $[H_2(g)]$, or both, increases the production of toluene. Since the number of gas molecules increases as the reaction proceeds from left to right, an increase in the volume of the system in the reaction vessel (decrease in pressure) will result in a shift to the right and increase the production of toluene. Since the reaction is endothermic from left to right, an increase in temperature will cause a shift to the right to minimize this change. Therefore, an increase in temperature will increase the production of toluene.

Student Textbook page 655

1. (a) $K_c = \dfrac{[SbCl_3(g)][Cl_2(g)]}{[SbCl_5(g)]}$

(b) $K_c = \dfrac{[N_2(g)][H_2O(g)]^2}{[H_2(g)]^2[NO(g)]^2}$

(c) $K_c = \dfrac{[H_2(g)]^4[CS_2(g)]}{[H_2S(g)]^2[CH_4(g)]}$

2. (a) $K_c = \dfrac{[N_2(g)][H_2(g)]^3}{[NH_3(g)]^2}$

(b) Equilibrium concentrations must be calculated from the information given in the problem.

$$[N_2(g)] = \frac{0.10 \text{ mol}}{0.50 \text{ L}} = 0.20 \text{ mol/L}$$

$$[H_2(g)] = \frac{0.30 \text{ mol}}{0.50 \text{ L}} = 0.60 \text{ mol/L}$$

$$[NH_3(g)] = \frac{0.80 \text{ mol}}{0.50 \text{ L}} = 1.6 \text{ mol/L}$$

Substituting into the expression

$$K_c = \frac{(0.20)(0.60)^3}{(1.6)^2} = 0.017$$

The equilibrium constant is 0.017.

3. The equilibrium expression is $K_c = \dfrac{[PCl_5(g)]}{[PCl_3(g)][Cl_2(g)]}$.

Equilibrium concentrations are given in the problem. Substituting into the expression

$$K_c = 49 = \frac{[PCl_5(g)]}{(0.035)(0.035)}$$

$[PCl_5] = 0.060$ mol/L

4. The equilibrium constant is small and [NO] is in the numerator of the equilibrium expression. Therefore, the equilibrium concentration of NO(g) at this temperature is very small.

5. The equilibrium constant is 14.5 and $[CH_3OH]$ is in the numerator of the equilibrium expression. Therefore, the formation of methanol is favourable at this temperature.

6. **(a)** right

(b) right

(c) no effect

(d) right

(e) left

7. **(a)** high temperature

(b) low temperature

(c) high temperature

(d) low temperature

8. **(a)** no effect

(b) reaction shifts to the right

(c) reaction shifts to the left

Student Textbook page 658

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

21. 30

22. $[I_2(g)] = [Cl_2(g)] = 0.015$ mol/L

$[ICl(g)] = 0.14$ mol/L

23. 4.7

24. $[HBr(g)] = 0.045$ mol/L

$[H_2(g)] = [Br_2(g)] = 2.9 \times 10^{-6}$ mol/L

Answers to Practice Problems 25–29

Student Textbook page 660

For full solutions to the practice problems, visit **www.albertachemistry.ca,** Online Learning Centre, Instructor Edition, Full Solutions.

25. $[CO(g)] = [Cl_2(g)] = 1.8 \times 10^{-4}$ mol/L

26. $[H_2(g)] = 9.4 \times 10^{-2}$ mol/L

27. $[CO_2(g)] = 0.60$ mol/L

$[CO(g)] = 1.1 \times 10^{-2}$ mol/L

$[O_2(g)] = 5.6 \times 10^{-3}$ mol/L

28. 6.0×10^{-5} mol/L

29. $[NO(g)] = 1.2 \times 10^{-5}$ mol/L

Student Textbook page 665

1. These are all homogeneous gas phase equilibrium reactions and the state (g) has been dropped in each expression for K_c, which is common practice.

(a) $K_c = \dfrac{[H_2]^4[CS_2]}{[H_2S]^2[CH_4]}$

(b) $K_c = \dfrac{[P_2O_3]^2}{[P_4][O_2]^3}$

(c) $K_c = \dfrac{[N_3S_4]^2[N_2S_2]^4}{[N_2]^7[S_8]^2}$

2. The ICE table is

$$2NH_3(g) \rightleftharpoons N_2(g) + 3H_2(g)$$

	$[NH_3(g)]$ (mol/L)	$[N_2(g)]$ (mol/L)	$[H_2(g)]$ (mol/L)
Initial	2.0	0	0
Change	$-2x$	$+x$	$+3x$
Equilibrium	1.6	0.20	0.60

$$K_c = \frac{[N_2][H_2]^3}{[NH_3]^2}$$

$$K_c = \frac{(0.20) \times (0.60)^3}{(1.6)^2}$$

$$K_c = 0.017$$

3. The ICE table is

$$SO_2(g) + NO_2(g) \rightleftharpoons NO(g) + SO_3(g)$$

	$[SO_2]$ (mol/L)	$[NO_2]$ (mol/L)	[NO] (mol/L)	$[SO_3]$ (mol/L)
Initial	0.36	0.36	0	0
Change	$-x$	$-x$	$+x$	$+x$
Equilibrium	$0.36 - x$	$0.36 - x$	x	x

$$K_c = \frac{[NO][SO_3]}{[SO_2][NO_2]}$$

$$K_c = \frac{x \times x}{(0.36 - x)(0.36 - x)}$$

This is a perfect square.

$$\sqrt{4.8} = \frac{x}{0.36 - x}$$

$$x = 0.247$$

$$[SO_3] = 0.247 \text{ mol/L}$$

In a 5.0 L container at equilibrium at the same temperature there will be

$0.247 \text{ mol/L} \times 5.0 \text{ L}$

$= 1.2 \text{ mol } SO_3(g).$

4. The ICE table is

$$PCl_3(g) + Cl_2(g) \rightleftharpoons PCl_5(g)$$

	$[PCl_3]$ (mol/L)	$[Cl_2]$ (mol/L)	$[PCl_5]$ (mol/L)
Initial	0.0938	0.0938	0
Change	$-x$	$-x$	$+x$
Equilibrium	0. 0938 – 0.059	0. 0938 – 0.059	0.059
	= 0.0348	= 0.0348	

$$K_c = \frac{[PCl_5]}{[PCl_3][Cl_2]}$$

$$K_c = \frac{0.059}{(0.0348)(0.0348)}$$

$$K_c = 49$$

5. (a) The ICE table is

$$2H_2O(g) \rightleftharpoons 2H_2(g) + O_2(g)$$

	$[H_2O]$ (mol/L)	$[H_2]$ (mol/L)	$[O_2]$ (mol/L)
Initial	0.055	0	0
Change	$-2x$	$+2x$	$+x$
Equilibrium	$0.055 - 2x$	$2x$	x
	≈ 0.055		

Students should be able to see, without performing the calculation, that

$$\frac{0.055}{7.3 \times 10^{-18}} >> 1000$$

Because x (and $2x$) is very small compared with 0.055, the approximation is valid.

$$K_c = \frac{[H_2]^2[O_2]}{[H_2O]^2}$$

$$7.3 \times 10^{-18} = \frac{(2x)^2(x)}{(0.055)^2}$$

$$x = 1.77 \times 10^{-7} \text{ mol/L}$$

$$[H_2] = 2 \times 1.77 \times 10^{-7} \text{ mol/L}$$

$$[H_2] = 3.5 \times 10^{-7} \text{ mol/L}$$

(b) This is not a practical method, because the equilibrium concentration of $H_2(g)$ is very small.

Student Textbook page 674

1. (a) The solubility of a gas dissolved in water is directly proportional to the partial pressure of this gas above the solution. Therefore, the solubility of a gas in water increases with an increase in the partial pressure of the gas above the solution. When the diver breathes air at greater pressure, larger concentrations of $O_2(aq)$ and $N_2(aq)$ appear in the blood. As the student textbook mentions, each of these increases in concentration of dissolved gas represents a potential hazard to the diver.

(b) The scuba diver must take into account the physiological effects of increased concentrations of $N_2(aq)$ (nitrogen narcosis) and $O_2(aq)$ (toxic). When ascending from depths greater than about 20 m, they must plan a stop, or stops, at prescribed depth for a certain period of time to allow the dissolved gases to leave the blood safely.

(c) Increased concentration of $N_2(aq)$ causes an effect called nitrogen narcosis. The effect is similar to consuming an alcoholic drink, and a diver might experience impaired judgment, confusion, and drowsiness. At depths of 90 m, nitrogen narcosis can lead to hallucinations and unconsciousness. Oxygen is toxic at increased concentration in the blood, and toxic levels can lead to a seizure.

2. The calcium carbonate shell will react with hot acidic water from hydrothermal vents.

3. (a) To increase the yield of hydrogen gas, the reaction should take place at low pressure.

(b) The reaction is endothermic. Therefore, to increase the yield of hydrogen gas the reaction should take place at a high temperature. This will increase the rate of the reaction.

(c) The catalyst has no effect on the position of equilibrium.

(d) $K_c = \dfrac{[H_2(g)]^4[CO_2(g)]}{[CH_4(g)][H_2O(g)]^2}$

(e) The pressure is raised to increase the yield of ammonia. The reaction is exothermic, so Le Châtelier's principle predicts that the forward reaction is favoured at low temperatures. However, this would slow the rate of reaction.

(f) $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g)$

(g) $K_c = \dfrac{[NH_3]^2}{[N_2][H_2]^3}$

(h) Substituting the equilibrium values into the expression for K_c

$$5.0 \times 10^{-2} = \frac{[NH_3]^2}{(1.2) \times (2.0)^3}$$

$$[NH_3] = 0.69 \text{ mol/L}$$

4. As the temperature increases, the value of K_c also increases. That is, as the temperature increases, the products are favoured. The reaction must be endothermic in the forward direction.

Student Textbook pages 676–677

Answers to Understanding Concepts Questions

1. The statement is almost certainly false. At equilibrium, the rate of change in opposite directions is equal, but the amounts are determined by the position of equilibrium.

2. Chemical equilibrium is a dynamic state because, at the molecular level, both the forward and the reverse reaction continue to take place.

3. In a chemical system at equilibrium, the concentrations of reactants and products are not in the same ratio as the coefficients in the chemical equation. The coefficients in a chemical equation describe the ratio in which substances react, not the concentrations at equilibrium.

4. **(a)** heterogeneous

   (b) homogeneous

   (c) heterogeneous

5. **(a)** $K_c = \dfrac{[CO_2(g)]^3[H_2O(g)]^4}{[C_3H_8(g)][O_2(g)]^5}$

   (b) $K_c = \dfrac{[NO(g)]^2[H_2O(g)]^2}{[N_2H_4(g)][O_2(g)]^2}$

   (c) $K_c = \dfrac{[Cu(NH_3)_4{}^{2+}(aq)]}{[Cu^{2+}(aq)][NH_3(aq)]^4}$

6. For a reaction that goes to completion, K_c is very large. The equilibrium expression has product terms in the numerator, and reactant terms in the denominator. For a reaction that goes to completion, there will be a very small concentration of reactant. Thus, K_c will be very large.

7. The factors that can affect the equilibrium of a reaction are: the concentration of reactants or products; the pressure of a gaseous system where the number of gas molecules changes during the reaction; and temperature. Of these factors, only temperature affects the value of K_c.

8. The reaction shifts to the right at higher temperature. Therefore, according to Le Châtelier's principle, the reaction is endothermic.

9. **(a)** Le Châtelier's principle predicts the equilibrium will shift to the right if dilute acid is added. The colour of the mixture should be more orange than before the addition of acid.

(b) If dilute base is added, it will remove $H_3O^+(aq)$ from the mixture. The equilibrium will shift to the left, and the mixture will become more yellow.

10. Changing the volume of the mixture has no effect in this case because there is the same number of gas molecules on each side of the equation. A decrease in temperature will shift the equilibrium to the right because the reaction is exothermic. Le Châtelier's principle predicts that increasing the temperature will shift the position of equilibrium towards the endothermic reaction, thus reducing the effect of the change.

11. An increase in pressure will result in a shift to the left where there are fewer gas molecules.

12. Pumping in a non-reacting gas has no effect on the equilibrium. This is because there is no change in the concentrations of the reacting gases.

13. Cousteau used a mix of helium and oxygen (Heliox) with relatively small $O_2(g)$ concentration to make sure his blood oxygen concentration did not reach toxic levels.

14. **(a)** Increasing the pressure of methane would shift the equilibrium to the right, increasing the concentration of hydrogen.

(b) Increasing the pressure of carbon dioxide would shift the equilibrium to the left, decreasing the concentration of hydrogen.

(c) Decreasing the pressure of water vapour would shift the equilibrium to the left, decreasing the concentration of hydrogen.

(d) The reaction is endothermic and raising the temperature would shift the equilibrium to the right, increasing the concentration of hydrogen.

(e) Adding a catalyst would have no effect on the concentration of hydrogen.

(f) Adding helium gas to the mixture would have no effect on the concentration of hydrogen.

(g) Transferring the mixture of gases to a container with greater volume would have the effect of reducing the pressure of the reacting mixture. The equilibrium would shift to the right, increasing the concentration of hydrogen.

Answers to Applying Concepts Questions

15. (a) $4HCl(g) + O_2(g) \rightleftharpoons 2H_2O(g) + 2Cl_2(g)$

(b) $$K_c = \frac{[H_2O(g)]^2[Cl_2(g)]^2}{[HCl(g)]^4[O_2(g)]}$$

$$= \frac{(7.8 \times 10^{-3})^2(3.6 \times 10^{-3})^2}{(2.7 \times 10^{-2})^4(8.6 \times 10^{-2})}$$

$$= 1.7 \times 10^{-2}$$

16. The chemical equation is $2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$

$$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$$

Substituting equilibrium concentrations from experiment 1

$$K_c = \frac{(3.50)^2}{(1.50)^2(1.25)} = 4.36$$

Substituting equilibrium concentrations from experiment 2

$$K_c = \frac{(0.260)^2}{(0.590)^2(0.0450)} = 4.32$$

These values are probably equal, within experimental error.

17. $4NH_3(g) + 5O_2(g) \rightleftharpoons 4NO(g) + 6H_2O(g)$

Since $[NO(g)] = [NH_3(g)]$, and these terms are raised to the same power, they divide out.

$$K_c = \frac{(2.0)^6}{(3.0)^5} = 0.26$$

18. $K_c = \dfrac{[CH_3OH(g)]}{[CO(g)][H_2(g)]^2}$

$$14.5 = \frac{[CH_3OH(g)]}{(0.25)(0.15)^2}$$

$$[CH_3OH(g)] = 0.082 \text{ mol/L}$$

19. The ICE table is

$$2SO_2(g) + O_2(g) \rightleftharpoons 2SO_3(g)$$

	$[SO_2]$ (mol/L)	$[O_2]$ (mol/L)	$[SO_3]$ (mol/L)
Initial	3.8	3.8	0
Change	$-2x$	$-x$	$+2x$
Equilibrium	$3.8 - 2x$	$3.8 - x$	$2x$
	≈ 3.8	≈ 3.8	

Check the assumption that x is negligible compared with 3.8

$$\frac{3.8}{3.6 \times 10^{-3}} = 1056$$

Since this is greater than 1000, the approximation is valid.

$$K_c = \frac{[SO_3(g)]^2}{[SO_2(g)]^2[O_2(g)]}$$

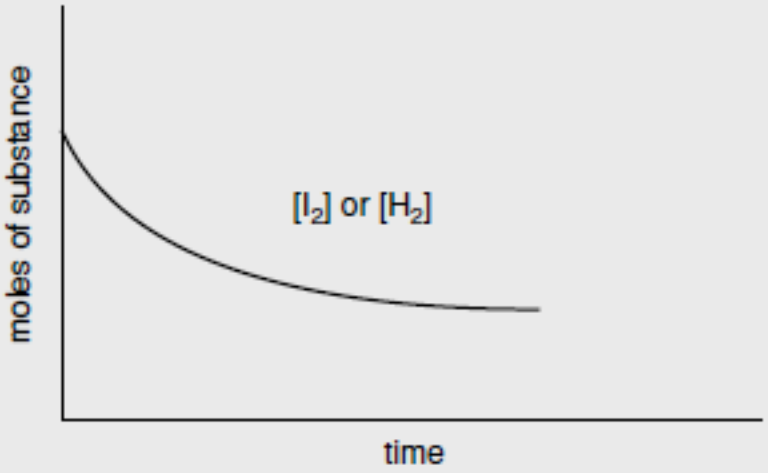
$$3.6 \times 10^{-3} = \frac{(2x)^2}{(3.8)^2(3.8)}$$

$$x = 0.222$$

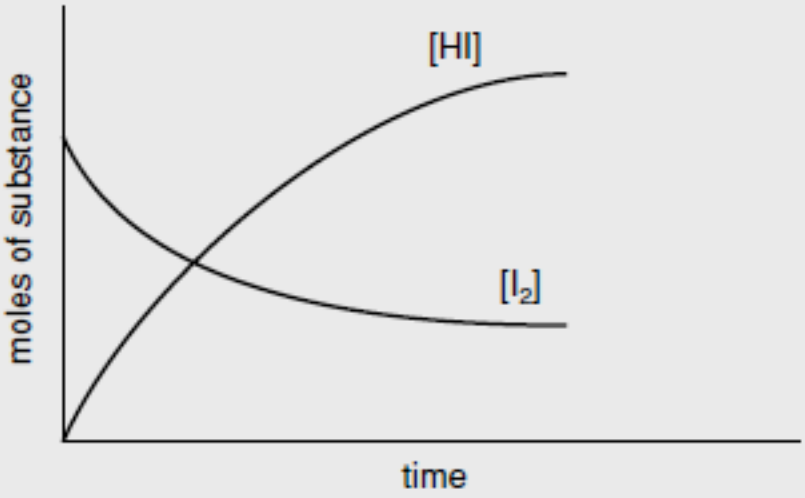
$$[SO_3(g)] = 2x = 0.44 \text{ mol/L}$$

Answers to Solving Problems Questions

20. (a)



(b). The graph of $[I_2(g)]$ and $[HI(g)]$ will be much different.



21. Carbon monoxide is dangerous because it is a highly poisonous, colourless and odourless gas. The gas forms in any situation where there is incomplete combustion caused by inadequate supply of air. The flyer should mention some of the common causes, such as operating a

gasoline engine in a closed garage, or a natural gas furnace with a blocked supply of air. The flyer should list the symptoms of CO(g) poisoning (headache and drowsiness), and suggest that homeowners check their furnaces for blockages (perhaps caused by nests) before the weather turns cold. The flyer should also mention that homeowners can easily install a CO(g) detector.

Answers to Making Connections Questions

22. The value for the equilibrium constant of a reaction is not related to the rate of the reaction.

23. Sulfide ions, $S^{2-}(aq)$, will combine with hydronium ions and remove them from solution. Consequently, the urine will become more basic and the pH will increase.

24. **(a)** The reaction conditions that favour increasing the yield of nitrogen monoxide are low pressure and low temperature.

(b) The catalyst has no effect on the position of equilibrium. A catalyst favours the production of NO(g) because it will increase the rate of the reaction.

(c) The reaction temperature is relatively high because, although the yield decreases at higher temperature, this is more than offset by an increase in the rate of the reaction.

(d) A relatively low pressure is used because this favours the production of NO(g). The cost of increasing the pressure in the system would not be justified.

(e) The reaction is exothermic, so cooling the gases forces the equilibrium towards more product. Also, it is probable that cooling the gases is a safety measure to prevent the reaction vessel from becoming dangerously overheated. The heat extracted would be used elsewhere in the plant.

(f) The nitrogen monoxide that is formed is recycled back to react with oxygen.

(g) Nitric acid has many uses including the manufacture of fertilizers, explosives, dyes, nylon, and pharmaceuticals.